

Chapter-15

Lesson-1

I - Systems at Equilibrium

A) The Concept of Equilibrium

- Equilibrium applies to the **extent** of a reaction, the concentration (amount) of product that has appeared given unlimited time.
- A system is at equilibrium when the rate of a reaction **is equal to** the rate of its reverse reaction.

$$\text{Rate}_{\text{forward reaction}} = \text{Rate}_{\text{reverse reaction}}$$

- A system is at equilibrium when the concentrations of its component parts (reactants and products) attain **constant** values.
- At the macroscopic level it may appear that the two opposing reactions have stopped once equilibrium has been attained. However, in most cases both reactions are still occurring but occurring at equal rates! This is called **dynamic equilibrium**.
- The **reaction quotient** (Q_c) of a system is a ratio of the current product concentrations multiplied together divided by the current reactant concentrations multiplied together, each raised to the power of their balancing coefficient.
- Only the concentrations of gaseous or aqueous constituents are included in the **reaction quotient** (Q_c) because they fill the entire volume of the system.
- For the reaction, $aA_{(g)} + bB_{(g)} \rightleftharpoons cC_{(g)} + dD_{(g)}$

The reaction quotient is, $Q_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$

← **products**
← **reactants**

- A system is at equilibrium when the **reaction quotient** (Q) of the system has reached a **constant** value.

B) The Equilibrium Constant (K_c)

- The **equilibrium expression** is a ratio of terms for a given reaction *at equilibrium* consisting of product the equilibrium concentrations multiplied together divided by the equilibrium reactant concentrations multiplied together, each raised to the power of their balancing coefficient.
- Only the concentrations of gaseous or aqueous constituents are included in the **equilibrium expression** because they fill the entire volume of the system.
- The **equilibrium constant** (K_c) is the numerical value obtained from the **equilibrium expression**.
- **A system is at equilibrium when $Q_c = K_c$!**

- For the reaction, $4 \text{HCl}_{(\text{aq})} + \text{O}_{2(\text{g})} \rightleftharpoons 2 \text{H}_2\text{O}_{(\text{l})} + 2 \text{Cl}_{2(\text{g})}$

The **equilibrium expression** is, $K_c = \frac{[\text{Cl}_2]^2}{[\text{HCl}]^4[\text{O}_2]}$

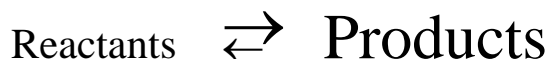
- For the reaction, $\text{CaCO}_{3(\text{s})} \rightleftharpoons \text{CaO}_{(\text{s})} + \text{CO}_{2(\text{g})}$

The **equilibrium expression** is, $K_c = [\text{CO}_2]$

- For the reaction, $\text{CO}_{(\text{g})} + 3 \text{H}_{2(\text{g})} \rightleftharpoons \text{CH}_{4(\text{g})} + \text{H}_2\text{O}_{(\text{l})}$

The **equilibrium expression** is, $K_c = \frac{[\text{CH}_4]}{[\text{CO}] [\text{H}_2]^3}$

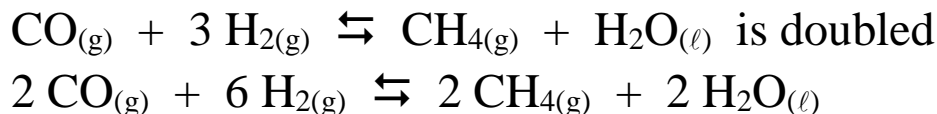
- The magnitude of the equilibrium constant (K_c) indicates the extent of the forward reaction and the relative equilibrium concentration of the products compared to the reactants.
- A large K_c value (greater than 1) usually indicates a large concentration of the products compared to the reactants.



- A small K_c value (less than 1) usually indicates a large concentration of the reactants compared to the products.



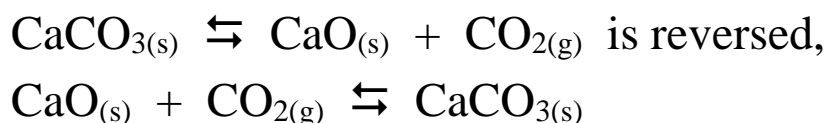
- ☛ When a reaction is doubled, the new equilibrium constant is the square of the original equilibrium constant.



The new equilibrium expression is, $K_c' = \frac{[\text{CH}_4]^2}{[\text{CO}]^2[\text{H}_2]^6}$

The new equilibrium constant K_c' equals K_c^2 .

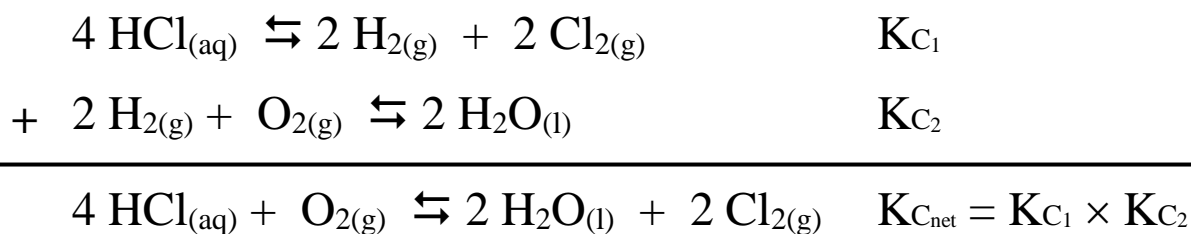
- ☛ When a reaction is reversed, the new equilibrium constant is the inverse of the original equilibrium constant.



The new equilibrium expression is, $K_c' = 1/[\text{CO}_2]$.

The new equilibrium constant K_c' equals $1/K_c$.

- ☛ When a reaction is a net reaction of individual steps, the net equilibrium constant for the net reaction is equal to the product of the equilibrium constants for the individual steps.



C) The Equilibrium Constant (K_p)

- When the reactants and products in a chemical reaction are gases, the **equilibrium expression** can also be represented as a ratio of the equilibrium partial pressures of the products multiplied together divided by the equilibrium partial pressures of the reactants multiplied together, each raised to the power of their balancing coefficient.
- The **equilibrium constant** (K_p) is the numerical value obtained from the **equilibrium expression**.

- For the reaction, $\text{N}_{2(\text{g})} + 3 \text{H}_{2(\text{g})} \rightleftharpoons 2 \text{NH}_{3(\text{g})}$

The **equilibrium expression** is,
$$K_p = \frac{(\text{P}_{\text{NH}_3})^2}{(\text{P}_{\text{N}_2})(\text{P}_{\text{H}_2})^3}$$

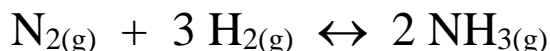
where P_{NH_3} is the partial pressure of NH_3 , P_{N_2} is the partial pressure of N_2 , and P_{H_2} is the partial pressure of H_2 .

- From the ideal gas law, $PV = nRT$, the partial pressure of a gaseous species in a system can be related to its molar concentration (M) or its mole fraction (X).

$$P_1 = \frac{n_1}{V}(RT) \quad \text{or} \quad P_1 = M_1(RT) \quad \text{or} \quad P_1 = X_1 P_t$$

- $K_p = K_c (RT)^{\Delta n_{\text{gas}}}$ where $R = 0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K}$
- $K_p = K_c \dots$ if $\Delta n_{\text{gas}} = 0$

Q1: Calculate the K_p for the following system at 300°C :



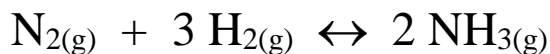
If the equilibrium constant (K_c) is 9.5

A1: $K_p = K_c(RT)^{\Delta n_{\text{gas}}}$

$$K_p = 9.5(0.08206 \times 573)^{-2}$$

$$K_p = 4.3 \times 10^{-3}$$

Q2: Calculate the K_c for the following system at 1000°C :



If the equilibrium constant (K_p) is 2.2×10^{-7} .

A2: $K_p = K_c(RT)^{\Delta n_{\text{gas}}}$

$$2.2 \times 10^{-7} = K_c(0.08206 \times 1273)^{-2}$$

$$K_c = 2.4 \times 10^{-3}$$

II – Calculating the Equilibrium Constant (K_c)

A) The Calculation of K_c Using Equilibrium Concentrations:

Q3: Calculate the K_c for the system: $2 \text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$ at 520°C if the equilibrium concentration of HI is 1.9 M, of H₂ is 0.24 M, and of I₂ is 0.24 M.

A3: Since all the given concentrations are equilibrium concentrations, they can be directly substituted into the equilibrium expression.

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{[0.24][0.24]}{[1.9]^2} = 0.016 \text{ or } 1.6 \times 10^{-2}$$

Q4: Calculate the K_c for the system: $2 \text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$ at 520°C if the initial concentration of HI is 0.200 M, of H₂ is 0.0 M, and of I₂ is 0.0 M and if the equilibrium concentrations of both H₂ and I₂ are 0.020 M.

A4: Since **NOT all** of the given concentrations are equilibrium concentrations, it is first necessary to determine all of the equilibrium concentrations and then to substitute them into the equilibrium expression. **The following chart is helpful!**

Step-1	Rx.	$2 \text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$		
	I	0.200 M	0.000 M	0.000 M
	C			
	E		0.020 M	0.020 M

Step-2	Rx.	$2 \text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$		
	I	0.200 M	0.000 M	0.000 M
	C		+0.020 M	+0.020 M
	E		0.020 M	0.020 M

Step-3	Rx.	$2 \text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$		
	I	0.200 M	0.000 M	0.000 M
	C	-0.040 M	+0.020 M	+0.020 M
	E		0.020 M	0.020 M

Step-4	Rx.	$2 \text{ HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$		
	I	0.200 M	0.000 M	0.000 M
	C	-0.040 M	+0.020 M	+0.020 M
	E	0.160 M	0.020 M	0.020 M

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{[0.020][0.020]}{[0.160]^2} = 0.016 \text{ or } 1.6 \times 10^{-2}$$

Q5: Calculate the K_c for the system: $2 \text{ HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$ at 520°C if the initial concentration of HI is 0.000 M, of H_2 is 0.200 M, and of I_2 is 0.200 M and if the equilibrium concentration of HI is 0.320 M.

A5: NOT all of the given concentrations are equilibrium concentrations

Rx.	$2 \text{ HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$		
I	0.000 M	0.200 M	0.200 M
C	+0.320 M	-0.160 M	-0.160 M
E	0.320 M	0.040 M	0.040 M

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{[0.040][0.040]}{[0.320]^2} = 0.016 \text{ or } 1.6 \times 10^{-2}$$

Q6: Calculate the K_c for the system: $2 \text{ HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$ at 520°C if the initial concentration of HI is 0.100 M, of H_2 is 0.100 M, and of I_2 is 0.100 M and if the equilibrium concentration of H_2 is 0.030 M.

A6: NOT all of the given concentrations are equilibrium concentrations

Rx. SYSTEM	$2 \text{ HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$		
Initial Conc.	0.100 M	0.100 M	0.100 M
Δ Conc.	+0.140 M	-0.070 M	-0.070 M
Equil. Conc.	0.240 M	0.030 M	0.030 M

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{[0.030][0.030]}{[0.240]^2} = 0.016 \text{ or } 1.6 \times 10^{-2}$$

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Lesson-2

III - Applications of the Equilibrium Constant

- The **reaction quotient** (Q_c) of a system is a number obtained by substituting the current reactant and product concentrations or partial pressures (Q_p) into an equilibrium expression.
- If the $Q_c = K_c$ (or $Q_p = K_p$), then the system is already at equilibrium.
- If the $Q_c < K_c$ (or $Q_p < K_p$), then the system has too little product and thus the system will “shift to the right” in order to achieve equilibrium.
- If the $Q_c > K_c$ (or $Q_p > K_p$), then the system has too much product and thus the system will “shift to the left” in order to achieve equilibrium.

Q1: If $K_c = 0.016$ for the system $2 \text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$ and it starts with all species at 0.10M, then in which direction will the system “shift” as it attains equilibrium?

A1: $Q_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{[0.10][0.10]}{[0.10]^2} = 1.0$

$Q_c > K_c$ there is too much product so the system will “shift to the left” to achieve equilibrium.

Q2: If $K_c = 0.36$ for the system $\text{N}_2\text{O}_{4(g)} \rightleftharpoons 2 \text{NO}_{2(g)}$ and it starts with 0.20 M N_2O_4 and 0.25 M NO_2 , then in which direction will the system “shift” as it attains equilibrium?

A2: $Q_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = \frac{[0.25]^2}{[0.20]} = 0.31$

$Q_c < K_c$ there is slightly too much reactant so the system will “shift to the right” to achieve equilibrium.

Q3: If $K_c = 0.0160$ for the system $2 \text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$ and if a container starts with 0.200M HI, then what will be the equilibrium concentrations for all species?

A3: The system will, *without a doubt*, “shift” to the right!

Rx. SYSTEM	$2 \text{HI}_{(g)} \rightleftharpoons$	$\text{H}_{2(g)} +$	$\text{I}_{2(g)}$
Initial Conc.	0.200 M	0.000 M	0.000 M
Δ Conc.	- 2X	+ X	+ X
Equil. Conc.	0.200 - 2X	X	X

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2}$$

$$0.0160 = \frac{[X][X]}{[0.200 - 2X]^2}$$

$$0.0160 = \frac{[X]^2}{[0.200 - 2X]^2}$$

This is a quadratic equation!
But there is a shortcut here!

$$0.126 = \frac{[X]}{[0.200 - 2X]}$$

Take the square root on both sides of the equation!

$$0.0252 - 0.252X = X$$

$$0.0252 = 1.252X$$

$$X = \boxed{0.0201 \text{ M}}$$

Rx. SYSTEM	$2 \text{HI}_{(g)} \rightleftharpoons$	$\text{H}_{2(g)} +$	$\text{I}_{2(g)}$
Equil. Conc.	0.160 M	0.0201 M	0.0201 M

$$\frac{[0.0201][0.0201]}{[0.160]^2} = 0.0158 \dots \text{ **CLOSE ENOUGH!**}$$

Q4: If $K_c = 8.3 \times 10^{-4}$ for the system $\text{COCl}_{2(g)} \rightleftharpoons \text{CO}_{(g)} + \text{Cl}_{2(g)}$ and if a 10.0 L container starts with 5.00 mol COCl_2 , then what will be the equilibrium concentrations for all species?

A4: The system will, without a doubt, "shift" to the right!

Rx. SYSTEM	$\text{COCl}_{2(g)} \rightleftharpoons$	$\text{CO}_{(g)}$	+	$\text{Cl}_{2(g)}$
Initial Conc.	0.500 M	0.000 M		0.000 M
Δ Conc.	- X	+ X		+ X
Equil. Conc.	0.500 - X	X		X

$$K_c = \frac{[\text{CO}][\text{Cl}_2]}{[\text{COCl}_2]}$$

$$8.3 \times 10^{-4} = \frac{[X][X]}{[0.500 - X]}$$

$$8.3 \times 10^{-4} = \frac{[X]^2}{[0.500 - X]}$$

This is a quadratic equation!
But there is a shortcut here? ... maybe!

$$8.3 \times 10^{-4} = \frac{[X]^2}{[0.500]}$$

Since K_c is small then the value of X may also be very small, so assume that $0.500 - X \approx 0.500$!

$$X^2 = 0.000415$$

$$X = \boxed{0.020 \text{ M}}$$

$$\frac{0.020}{0.500} = 4\%$$

Since 0.020 M is less than 5% of the original 0.500M, the assumption is justified!

Rx. SYSTEM	$\text{COCl}_{2(g)} \rightleftharpoons$	$\text{CO}_{(g)}$	+	$\text{Cl}_{2(g)}$
Equil. Conc.	0.480 M	0.020 M		0.020 M

$$\frac{[0.020]^2}{[0.480]} = 8.3 \times 10^{-4} \dots \text{YEAH!}$$

Q5: If $K_c = 8.3 \times 10^{-4}$ for the system $\text{COCl}_{2(g)} \rightleftharpoons \text{CO}_{(g)} + \text{Cl}_{2(g)}$ and if a container is initially injected with 0.150 M COCl_2 , then what will be the equilibrium concentrations for all species?

A5: The system will, without a doubt, "shift" to the right!

Rx. SYSTEM	$\text{COCl}_{2(g)} \rightleftharpoons$	$\text{CO}_{(g)}$	+	$\text{Cl}_{2(g)}$
Initial Conc.	0.150 M	0.000 M		0.000 M
Δ Conc.	- X	+ X		+ X
Equil. Conc.	0.150 - X	X		X

$$K_c = \frac{[\text{CO}][\text{Cl}_2]}{[\text{COCl}_2]}$$

$$8.3 \times 10^{-4} = \frac{[X]^2}{[0.150 - X]}$$

This is a quadratic equation!
The shortcut may work again!

$$8.3 \times 10^{-4} = \frac{[X]^2}{[0.150]}$$

Since K_c is small then the value of X may be small too, so assume that $0.150 - X \approx 0.150$!

$$X = 0.011 \text{ M}$$

$$\frac{0.011}{0.150} = 7.3\%$$

Since 0.011 M is more than 5% of the original 0.150M, the assumption NOT is justified!

$$8.3 \times 10^{-4} = \frac{[X]^2}{[0.150 - 0.011]}$$

Since 7.3% is not way above 5%, the method of successive approximations will work!

$$X = 0.0107 \text{ M (1}^{\text{st}})$$

$$8.3 \times 10^{-4} = \frac{[X]^2}{[0.150 - 0.0107]}$$

$$\rightarrow X = 0.0108 \text{ M (2}^{\text{nd}})$$

$$8.3 \times 10^{-4} = \frac{[X]^2}{[0.150 - 0.0108]}$$

$$\rightarrow X = 0.0107 \text{ M (3}^{\text{rd}})$$

Rx. SYSTEM	$\text{COCl}_{2(g)} \rightleftharpoons$	$\text{CO}_{(g)}$	+	$\text{Cl}_{2(g)}$
Equil. Conc.	0.13925 M	0.01075 M		0.01075 M

$$\frac{[0.01075]^2}{[0.13925]} = 8.3 \times 10^{-4} \dots \text{YEAH!}$$

Q6: If the $K_c = 0.36$ for the system $\text{N}_2\text{O}_{4(g)} \rightleftharpoons 2 \text{NO}_{2(g)}$ and if a container started with 0.100 M $\text{N}_2\text{O}_{4(g)}$, then what will be the equilibrium concentrations for all species in the container?

A6: The system will, without a doubt, "shift" to the right!

Rx. SYSTEM	$\text{N}_2\text{O}_{4(g)} \rightleftharpoons 2 \text{NO}_{2(g)}$	
Initial Conc.	0.100 M	0.000 M
Δ Conc.	- X	+ 2X
Equil. Conc.	0.100 - X	2X

$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$

$$0.36 = \frac{[2X]^2}{[0.100 - X]}$$

$$0.36 = \frac{4X^2}{0.100 - X} \quad \text{Let's assume that "- X" is small.}$$

$$0.36 = \frac{4X^2}{0.100}$$

$$X = 0.095 \text{ M} \quad \text{YIKES! This is 95\% of the original... way too big!}$$

$$0.036 - 0.36X = 4X^2$$

The quadratic formula is needed!

$$4X^2 + 0.36X - 0.036 = 0$$

$$X = \frac{-b \pm [b^2 - 4ac]^{1/2}}{2a}$$

$$X^2 + 0.09X - 0.009 = 0$$

$$X = \frac{-0.09 \pm [(0.09)^2 - (4 \times 1 \times -0.009)]^{1/2}}{2 \times 1}$$

$$X = 0.060 \text{ M} \quad \text{or} \quad -0.15 \text{ M (this root is ridiculous!)}$$

Rx. SYSTEM	$\text{N}_2\text{O}_{4(g)} \rightleftharpoons 2 \text{NO}_{2(g)}$	
Equil. Conc.	0.040 M	0.120 M

$$\frac{[2X]^2}{[0.100 - X]} = \frac{[2 \times 0.060]^2}{[0.100 - 0.060]} = 0.36 \dots \text{YEAH!}$$

Q7: If the $K_c = 4.0 \times 10^{-11}$ for the system $2 \text{SO}_{3(g)} \rightleftharpoons 2 \text{SO}_{2(g)} + \text{O}_{2(g)}$ and if a container started with 0.100 M $\text{SO}_{3(g)}$, then what will be the equilibrium concentrations for all species in the container?

A7: The system will, without a doubt, "shift" to the right!

	$2 \text{SO}_{3(g)} \rightleftharpoons 2 \text{SO}_{2(g)} + \text{O}_{2(g)}$		
Initial Conc.	0.100 M	0.000 M	0.000 M
Δ Conc.	- 2X	+ 2X	+ X
Equil. Conc.	0.100 - 2X	2X	X

$$K_p = \frac{[\text{SO}_2]^2[\text{O}_2]}{[\text{SO}_3]^2}$$

$$4.0 \times 10^{-11} = \frac{[2X]^2[X]}{[0.100 - 2X]^2} \quad \text{Let's assume that "- 2X" is small.}$$

$$4.0 \times 10^{-11} = \frac{4X^3}{0.0100}$$

$$X = 4.6 \times 10^{-5} \text{ M}$$

Rx. SYSTEM	$2 \text{SO}_{3(g)} \rightleftharpoons 2 \text{SO}_{2(g)} + \text{O}_{2(g)}$		
Equil. Conc.	0.0999 M	$9.2 \times 10^{-5} \text{ M}$	$4.6 \times 10^{-5} \text{ M}$

$$\frac{[2X]^2[X]}{[0.100 - 2X]^2} = \frac{[9.2 \times 10^{-5}]^2[4.6 \times 10^{-5}]}{[0.100 - 9.2 \times 10^{-5}]^2} = 3.9 \times 10^{-11}$$

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Lesson-3

IV - Le Chatelier's Principle

- If a system at equilibrium is subjected to a stress (a change in condition) that results in a loss of equilibrium, the system will "shift" in the direction that will partially relieve the stress. This is known as a *Le Chatelier's principle*.

A) Adding or Removing a Gaseous Species:

Q1: If the system $2 \text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$ is at equilibrium and some $\text{H}_{2(g)}$ were added, then the system will shift to the _____ as it regains equilibrium. Once equilibrium has been regained, the $[\text{HI}]$ will be _____ than before, the $[\text{H}_2]$ will be _____ than before, and the $[\text{I}_2]$ will be _____ than before.

A1: *LEFT, MORE, MORE, LESS*

Q2: If the system $2 \text{NH}_{3(g)} \rightleftharpoons \text{N}_{2(g)} + 3 \text{H}_{2(g)}$ is at equilibrium and some $\text{N}_{2(g)}$ were removed, then the system will shift to the _____ as it regains equilibrium. Once equilibrium has been regained, the $[\text{NH}_3]$ will be _____ than before, the $[\text{N}_2]$ will be _____ than before, and the $[\text{H}_2]$ will be _____ than before.

A2: *RIGHT, LESS, LESS, MORE*

Q3: If the system $\text{CO}_{2(s)} \rightleftharpoons \text{CO}_{2(g)}$ is at equilibrium and some $\text{CO}_{2(s)}$ were removed, then the system would... _____

A3: *NOT CHANGE BECAUSE AS LONG AS SOME SOLID WAS LEFT THE EQUILIBRIUM POINT WOULD NOT CHANGE!*

B) Changing Volume or Pressure:

Q4: If the system $\text{N}_2\text{O}_{4(g)} \rightleftharpoons 2 \text{NO}_{2(g)}$ is at equilibrium and the volume of the system is increased, then the system will "shift" to the _____ as it regains equilibrium. Once equilibrium has been regained, the $[\text{N}_2\text{O}_4]$ will be _____ than before and the $[\text{NO}_2]$ will be _____ than before.

A4: *RIGHT, LESS, MORE*

Q5: If the system $2 \text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$ is at equilibrium and the pressure on the system were increased, then the system will _____ and the $[\text{HI}]$ will be _____ as before, the $[\text{H}_2]$ will be _____ as before, and the $[\text{I}_2]$ will be _____ as before.

A5: *NOT CHANGE, THE SAME, THE SAME, THE SAME*

C) Changing Temperature:

Q6: If the system $2 \text{SO}_{3(g)} + 198.2 \text{ kJ} \rightleftharpoons 2 \text{SO}_{2(g)} + \text{O}_{2(g)}$ is at equilibrium and the temperature of the system is increased, then the system will "shift" to the _____ as equilibrium is restored. Once equilibrium has been regained, the $[\text{SO}_3]$ will be _____ than before, the $[\text{SO}_2]$ will be _____ than before, and the $[\text{O}_2]$ will be _____ than before.

A6: *RIGHT, LESS, MORE, MORE*

Q7: If the system $\text{N}_{2(g)} + 3 \text{H}_{2(g)} \rightleftharpoons 2 \text{NH}_{3(g)} + 92.4 \text{ kJ}$ is at equilibrium and the temperature is increased, then the system will shift to the _____ as it regains equilibrium. Once equilibrium has been regained, the $[\text{NH}_3]$ will be _____ than before, the $[\text{N}_2]$ will be _____ than before, and the $[\text{H}_2]$ will be _____ than before.

A7: *LEFT, LESS, MORE, MORE*

Q8: If some $\text{HI}_{(\text{g})}$ were added to the equilibrium system below such that the concentration of $\text{HI}_{(\text{g})}$ suddenly increased by 0.032 M, then what would be the new equilibrium concentrations of each species in the system once equilibrium is regained? ($K_c = 0.016$)

Rx. SYSTEM	$2 \text{HI}_{(\text{g})} \rightleftharpoons \text{H}_{2(\text{g})} + \text{I}_{2(\text{g})}$
1 st Equil. Conc.	0.160 M 0.020 M 0.020 M

A8: Adding more HI will shift equilibrium to the right?

Rx. SYSTEM	$2 \text{HI}_{(\text{g})} \rightleftharpoons \text{H}_{2(\text{g})} + \text{I}_{2(\text{g})}$
1 st Equil. Conc.	0.160 M 0.020 M 0.020 M
Stress	+ 0.032
Δ Conc.	- 2X +X +X
2 nd Equil. Conc.	0.192 - 2X 0.020 + X 0.020 + X

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2}$$

$$0.016 = \frac{[0.020 + X]^2}{[0.192 - 2X]^2}$$

$$0.13 = \frac{[0.020 + X]}{[0.192 - 2X]}$$

$$X = \boxed{0.004 \text{ M}}$$

Rx. SYSTEM	$2 \text{HI}_{(\text{g})} \rightleftharpoons \text{H}_{2(\text{g})} + \text{I}_{2(\text{g})}$
1 st Equil. Conc.	0.160 M 0.020 M 0.020 M
2 nd Equil. Conc.	0.184 M 0.024 M 0.024 M

Equilibrium did indeed shift to the right. Notice that the new amount of $\text{HI}_{(\text{g})}$ is greater than before. The equilibrium “shift” reduced the amount of added $\text{HI}_{(\text{g})}$.

D) The Calculation of K_{eq} Using the Van't Hoff Equation:

- The van't Hoff equation is very similar to the Clausius-Clapeyron equation and the Arrhenius equation.

$$\ln \frac{K_2}{K_1} = - \frac{\Delta H_R^\circ}{R} \left[\frac{1}{T_2} - \frac{1}{T_1} \right]$$

Where K_1 is the equilibrium constant for the system at T_1 (in K)
and K_2 is the equilibrium constant for the system at T_2 (in K)
and ΔH_R (in J/mol) is the heat of reaction of the forward reaction.

Q10: Calculate the K_c for the system: $2 \text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$ at 25.0°C
if the K_c for the system at 520°C is 0.016 and if the ΔH_R° for the
decomposition of $\text{HI}_{(g)}$ is $+46.2 \text{ kJ/mol}$.

A10: $\ln \frac{K_2}{0.016} = \frac{(-46200)}{8.31} \left[\frac{1}{298} - \frac{1}{793} \right]$

$$\ln K_2 - \ln 0.016 = -11.6$$

$$\ln K_2 + 4.1 = -11.6$$

$$\ln K_2 = -15.7$$

$$K_2 = \boxed{1.52 \times 10^{-7}} \text{ At } 25.0^\circ\text{C, the reaction does not proceed very far!}$$

Q11: Calculate the K_c for the system: $\text{N}_2\text{O}_{4(g)} \rightleftharpoons 2 \text{NO}_{2(g)}$ at 100.0°C
if the K_c for the system at 25.0°C is 0.0032 and if the ΔH_R° for
the decomposition of $\text{N}_2\text{O}_{4(g)}$ is $+58.2 \text{ kJ/mol}$.

A11: $\ln \frac{K_2}{0.0032} = \frac{(-58200)}{8.31} \left[\frac{1}{373} - \frac{1}{298} \right]$

$$\ln K_2 - \ln 0.0032 = +4.73$$

$$\ln K_2 + 5.74 = +4.73$$

$$\ln K_2 = -1.01$$

$$K_2 = \boxed{0.36} \text{ At } 100.0^\circ\text{C, the reaction proceeds somewhat to the right!}$$